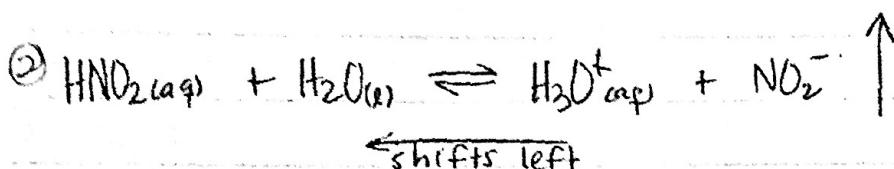
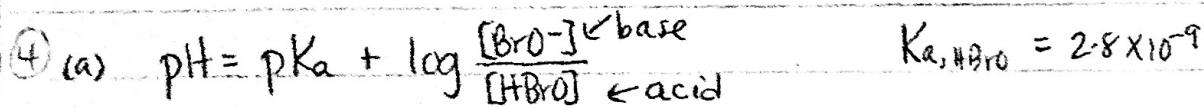
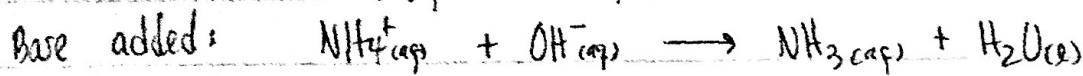
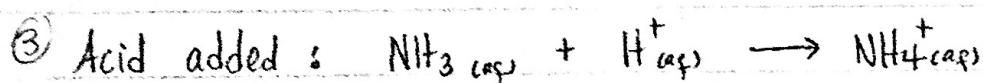


① (a) acidic  
(b) basic  
(c) neutral

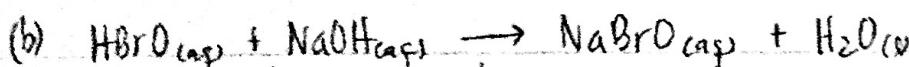
(d) acidic  
(e) neutral ( $K_a = K_b$ )



$\text{HNO}_2$  will dissociate less in  $\text{NaNO}_2$  because the presence of more  $\text{NO}_2^-$  in solution will shift  $\text{HNO}_2$ 's equilibrium left, which will result in less ionization.



$$\text{pH} = -\log(2.8 \times 10^{-9}) + \log\left(\frac{0.18}{0.15}\right) = 8.63$$



before Addition	0.15 mol	0	0.18 mol
Addition	—	0.01 mol	—
After Addition	0.14 mol	0	0.19 mol

$$\text{pH} = -\log(2.8 \times 10^{-5}) + \log\left(\frac{0.19}{0.14}\right) = \underline{\underline{8.109}}$$



	before Addition	After Addition	After Addition	After Addition
Addition	0.18 mol	—	0.01 mol	—
After Addition	0.17 mol	0	0	0.16 mol

(2)

$$pH = -\log(2.8 \times 10^{-5}) + \log \left( \frac{0.17}{0.16} \right) \stackrel{\text{base}}{\underset{\text{acid}}{\cancel{}}} = 8.58$$

$$\textcircled{5} \quad pH = pK_a + \log \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]} \stackrel{\text{base}}{\underset{\text{acid}}{\cancel{}}} \quad K_a, \text{H}_2\text{C}_2\text{H}_3\text{O}_2 = 1.8 \times 10^{-5}$$

$$5.00 = 4.74 + \log \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{0.150}$$

$$10^{0.26} = \frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{0.150} \Rightarrow [\text{NaC}_2\text{H}_3\text{O}_2] = (0.150)(10^{0.26}) = 0.27 \text{ M}$$

$$\text{mol NaC}_2\text{H}_3\text{O}_2 = (0.27 \text{ M})(1.00 \text{ L}) = 0.27 \text{ mol NaC}_2\text{H}_3\text{O}_2$$

$$0.27 \text{ mol NaC}_2\text{H}_3\text{O}_2 \left( \frac{82.04 \text{ g NaC}_2\text{H}_3\text{O}_2}{1 \text{ mol NaC}_2\text{H}_3\text{O}_2} \right) = 22 \text{ g NaC}_2\text{H}_3\text{O}_2$$

$$\textcircled{7} \quad \text{Equivalence Point: } C_a V_a = C_b V_b \Rightarrow V_a = \frac{C_b V_b}{C_a} = \frac{(0.240 \text{ M})(25.0 \text{ mL})}{0.200 \text{ M}}$$

Strong base - Strong acid  
titration

$$V_a = 30.0 \text{ mL} \quad \text{Equivalence Point}$$

(a) 0.0 mL

$$[\text{OH}^-] = [\text{NaOH}] = 0.240 \text{ M}$$

$$pOH = -\log(0.240) = 0.620$$

$$pH = 14.00 - 0.62 = 13.38$$



Excess NaOH

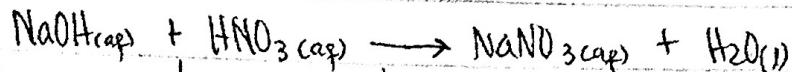
BA	6.00 mmol	-	-	-
A	-	3.00 mmol	-	-
AA	3.00 mmol	0	3.00 mmol	neutral salt.

$$[\text{OH}^-] = \frac{3.00 \text{ mmol}}{40.0 \text{ mL}} = 0.0750 \text{ M}$$

$$pOH = -\log(0.0750) = 1.125 \Rightarrow pH = 14.00 - 1.125 = 12.88$$

(3)

(c) 25.0 mL



BA	6.00 mmol	-	-	-
A	-	5.00 mmol	-	-
AA	1.00 mmol	0	5.00 mmol	Neutral salt

$$[\text{OH}^-] = \frac{1.00 \text{ mmol}}{50.0 \text{ mL}} = 0.0200 \text{ M}$$

$$\text{pOH} = -\log(0.0200) = 1.699 \Rightarrow \text{pH} = 14.00 - 1.699 = \underline{\underline{12.30}}$$

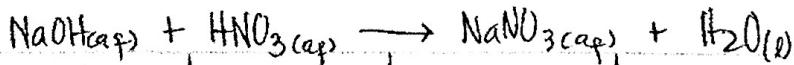
(d) 30.0 mL



BA	6.00 mmol	-	-	-
A	-	6.00 mmol	-	-
AA	0	0	6.00 mmol	Neutral salt

At equivalence point of a strong base-strong acid titration, a neutral salt is present in solution (no excess of strong acid or base present) so pH is 7.00.

(e) 45.0 mL



BA	6.00 mmol	-	-	-
A	-	9.00 mmol	-	-
AA	0	3.00 mmol	6.00 mmol	Neutral salt

$$[\text{H}_3\text{O}^+] = \frac{3.00 \text{ mmol}}{70.0 \text{ mL}} = 0.0429 \text{ M}$$

$$\text{pH} = -\log(0.0429) = \underline{\underline{1.368}}$$

Excess NaOH

Equivalence point

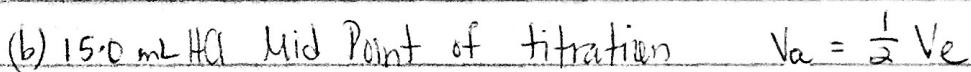
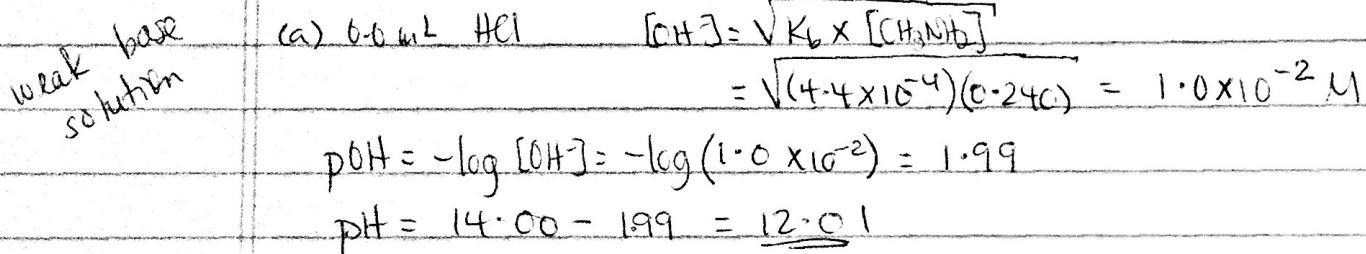
Excess HNO<sub>3</sub>

(4)

- ⑥ (a) weak acid / strong base titration pH is above 7.  
 (b) strong acid / strong base titration pH is 7.  
 (c) weak base / strong acid titration pH is below 7.

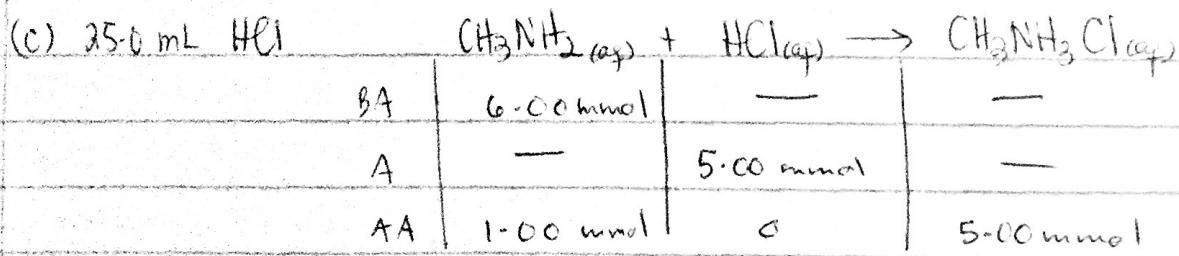
⑧ Equivalence Volume  $C_a V_a = C_b V_b \Rightarrow V_a = \frac{C_b V_b}{C_a}$   
 $= \frac{(25.0 \text{ mL})(0.240 \text{ M})}{0.200 \text{ M}}$

$V_a = 30.0 \text{ mL}$



Buffer region  
 $K_a \text{ of } CH_3NH_3^+ \Rightarrow K_a = \frac{K_w}{K_b} = \frac{1.0 \times 10^{-14}}{4.4 \times 10^{-4}} = 2.3 \times 10^{-11}$

At midpoint,  $pH = pK_a = -\log(2.3 \times 10^{-11}) = 10.64$



$pH = pK_a + \log \frac{[base]}{[acid]} = -\log(2.3 \times 10^{-11}) + \log \frac{(1.00)}{(5.00)} = 9.94$

(5)

Equivalence  
Point

(d) 30.0 mL HCl



BT	6.00 mmol	—	—
A	—	6.00 mmol	—
AA	0	0	6.00 mmol

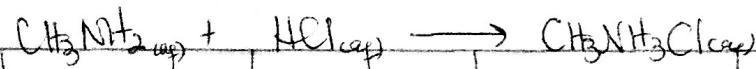
$$[\text{CH}_3\text{NH}_3^+] = \frac{6.00 \text{ mmol}}{55.0 \text{ mL}} = 0.109 \text{ M}$$

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \times [\text{CH}_3\text{NH}_3^+]} = \sqrt{(2.3 \times 10^{-1}) (0.109)} = 1.6 \times 10^{-4} \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (1.6 \times 10^{-4}) = \underline{\underline{5.80}}$$

Excess  
HCl

(e) 45.0 mL HCl



BT	6.00 mmol	—	—
A	—	9.00 mmol	—
AA	0	3.00 mmol	6.00 mmol

$$[\text{H}_3\text{O}^+] = [\text{HCl}] = \frac{3.00 \text{ mmol}}{70.0 \text{ mL}} = 0.0429 \text{ M}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log (0.0429) = \underline{\underline{1.368}}$$

$$\textcircled{3} \text{ Equivalence Point: } C_a V_a = C_b V_b \Rightarrow V_b = \frac{C_a V_a}{C_b} = \frac{(0.200 \text{ M})(25.0 \text{ mL})}{0.250 \text{ M}}$$

weak acid -  
strong base titration

$$V_{eq} = 20.0 \text{ mL}$$

Equivalence  
point.

Weak Acid  
solution

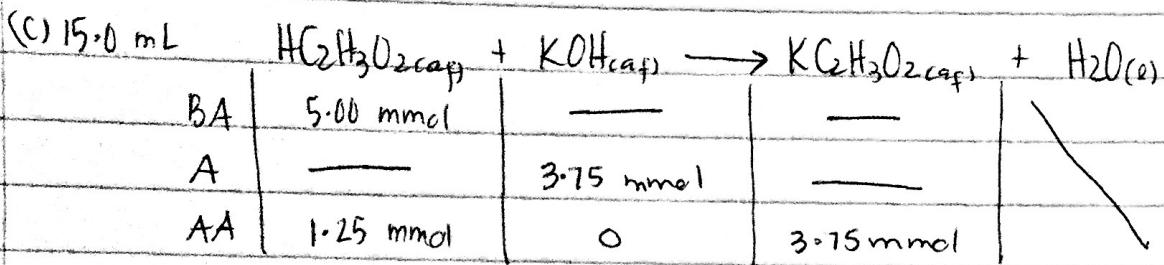
(a) 0.0 mL

$$[\text{H}_3\text{O}^+] = \sqrt{K_a \times [\text{H}_2\text{O}_2]} = \sqrt{(1.8 \times 10^{-5})(0.200)} = 1.9 \times 10^{-3} \text{ M}$$

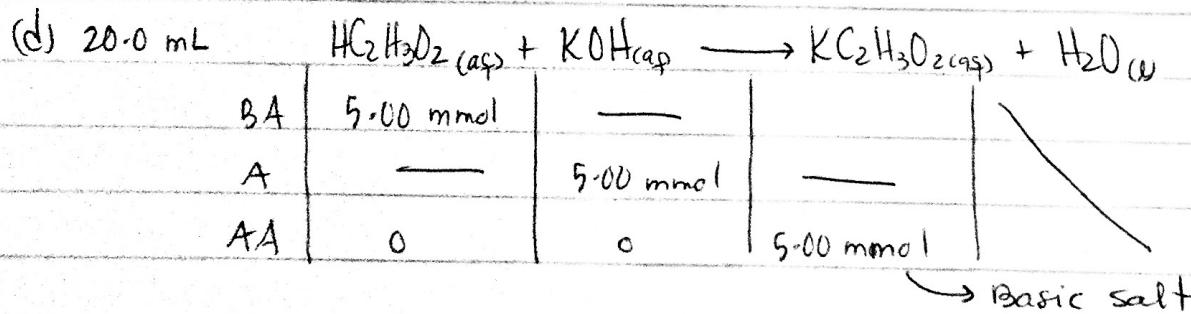
$$\text{pH} = -\log (1.9 \times 10^{-3}) = \underline{\underline{2.72}}$$

(6)

Buffer Region

(b) 10.0 mL  $V_b = \frac{1}{2} V_{eq}$  Mid Point of titrationAt midpoint  $pH = pK_a \Rightarrow pH = -\log(1.8 \times 10^{-5}) = 4.74$ 

$$pH = -\log(1.8 \times 10^{-5}) + \log\left(\frac{3.75}{1.25}\right) = 5.22$$

Equivalence point  
weak base solution

$$[\text{C}_2\text{H}_3\text{O}_2^-] = \frac{5.00 \text{ mmol}}{45.0 \text{ mL}} = 0.111 \text{ M} \quad K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

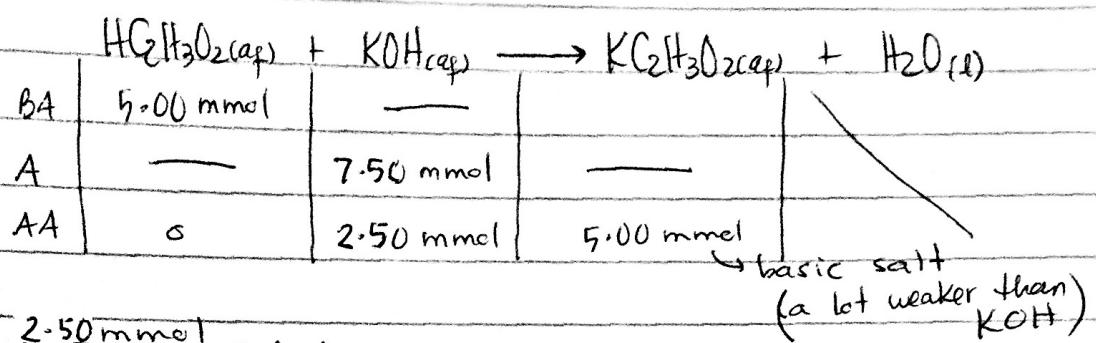
$$[\text{OH}^-] = \sqrt{K_b \times [\text{C}_2\text{H}_3\text{O}_2^-]} = \sqrt{(0.111)(5.6 \times 10^{-10})} = 7.9 \times 10^{-6} \text{ M}$$

$$\text{pOH} = -\log(7.9 \times 10^{-6}) = 5.10$$

$$\text{pH} = 14.00 - 5.10 = 8.90$$

(7)

(e) 30.0 mL



$$[\text{OH}^-] = \frac{2.50 \text{ mmol}}{55.0 \text{ mL}} = 0.0455 \text{ M}$$

$$\text{pOH} = -\log(0.0455) = 1.342$$

$$\text{pH} = 14.00 - 1.342 = \underline{12.66}$$

⑨ Plot titration curve  $\Rightarrow$  pH vs. mL NaOH added

At midpoint ( $V_{\text{NaOH}} = \frac{1}{2} V_{\text{equivalence}}$ ), pH = pKa

Equivalence Volume  $\approx 22.4 \text{ mL}$ ; midpoint volume = 11.2 mL

$$\text{pH at midpoint} \approx 4.50 = \text{pKa} \Rightarrow K_a = 10^{-\text{pKa}} = 10^{-4.50} = \underline{\underline{3.2 \times 10^{-5}}}$$

plot attached

### Titration of a monoprotic acid with a NaOH solution

